Precipitation Reactions

Many common chemical reactions involve ions and take place in aqueous solution. Examples include almost all single and double replacement reactions. A primary reason is that ions can react very easily when dissolved, but not under most other conditions.

In this lab, we will study one particular type of double replacement reaction: precipitation reactions. Precipitates are really just ionic compounds that have very low solubilities in water. Because certain combinations of ions cannot dissolve much, if too many of those ions are present, they instead form small particles of the solid chemical that settle out of solution.

How can precipitates form when we mix two solutions together? Before they are combined, each solution would contain dissolved ions only. But when two different solutions are mixed together, new ion combinations are possible. If the new combination is not very soluble, it will precipitate out. Consider the following classic example:

Each of the following compounds dissolve easily in water: NaCl and $AgNO_3$. This implies that in a solution of sodium chloride, there is NO solid NaCl... there are only Na^+ ions and Cl^- ions, thus they are aqueous, or "aq". Similarly with a solution of silver nitrate, $AgNO_3$. Once in solution, there are two possible NEW ion pairs: (1) Ag^+ with Cl^- and (2) Na^+ with NO_3^-. It turns out that the Na^+ and NO_3^- ions do NOT react... this is because they form a soluble compound (aq) thus they stay dissolved. But the Ag^+ and Cl^- ions react differently. AgCl is an insoluble * compound, one that does not dissolve. The ions cannot dissolve together in water... and they don't. They combine to make solid AgCl which precipitates out of solution.

* because all ionic compounds dissolve at least a little, a better definition for "insoluble" is a compound that has a very low solubility.

This information can be written as a "regular" equation showing the reaction between silver nitrate and sodium chloride. The balanced reaction would look like this:

$$AgNO_3\,(aq) + NaCl\,(aq) \dashrightarrow AgCl\,(s) + NaNO_3\,(aq)$$

In addition, the example mentioned above can be written as an ionic equation. Ionic equations give slightly different information than regular equations... they show chemicals which exist as ions in solutions as those ions, with the symbol (aq) to show they are dissolved. There are two types:

Complete ionic equation:

$$Ag^+\,(aq) + NO_3^-\,(aq) + Na^+(aq) + Cl^-\,(aq) \rightarrow AgCl\,(s) + NO_3^-\,(aq) + Na^+\,(aq)$$

Net ionic equation: $Ag^+\,(aq) + Cl^-\,(aq) \rightarrow AgCl\,(s)$

Note the following:

- In each equation the precipitate is a solid; all other particles are ions
- The complete ionic equation shows all ions in solution in their correct ratios.
- Certain ions do not actually react. in this case Na^+ and NO_3^-. These ions are called spectator ions. because they are just "looking on."
- The net ionic equation does not show the spectator ions; it only shows the ions involved in the precipitation itself; like all equations the coefficients are reduced to lowest ratios
- When ionic equations are balanced, the number of each type of atom on each side is balanced, and the total net charge of the ions on each side is balanced.

Objectives:

1) To observe double replacement reactions.
2) To practice writing complete ionic and net ionic equations and to identify spectator ions.
3) To determine the likelihood of ions being present in an unknown chemical.

Materials:

Plastic sheet or spotting plate for dropping chemicals
dropper bottles with chemicals listed below

Chemicals: (approximately 0.2 molar solutions)

Set A: Set B:

#	name	formula	#	name	formula
1	barium nitrate		1	magnesium chloride	
2		Na_2SO_4	2		$NaOH$
3	aluminum sulfate		3	potassium chloride	
4		$MgCl_2$	4		$BaCl_2$
5	magnesium nitrate		5	sodium sulfate	
6		$AlCl_3$	6		$MgSO_4$

Set C:

#	name	formula
1	potassium chromate	
2		$Al_2(SO_4)_3$
3	barium chloride	
4		Na_2CrO_4
5	silver nitrate	
6		$Mg(NO_3)_2$

Pre-lab (complete 1-5 on a separate sheet)

1. Fill in the missing chemical formulas or names in the chemical sets above.
2. Make a data table for this lab: hint: think about a grid in two dimensions. You may be graded on how quickly and easily I can interpret your work. If you design your experiment well, you can place it under your plastic sheet to guide your experiment (see Procedure). Make your table big enough so that if you add several drops of chemicals to each spot they won't run together. Even if you write the names of the chemicals, also use numbers because I will use the number code in checking your results.
3. Explain the difference between a complete ionic and a net ionic equation.
4. Write a complete ionic equation for the reaction of potassium sulfide and iron (III) iodide. The precipitate is iron (III) sulfide. Hint: first balance the regular reaction, then turn the aqueous formulas into ions!

Procedure

1. Test each set of chemicals for precipitates by adding one or two drops of each to a plastic sheet. Test all combinations (but remember... if you test "2 and 1", you do not need to test "1 and 2") Note... most precipitates form immediately, but hydroxides may take a minute or so
   - to save time, mix all combinations first, then check for precipitates
   - a piece of paper under your plastic sheet may guide where you place each spot
   - white precipitates are easiest to see on a dark background like your desk
   - your recorded data may be as simple as a "+" (there is a precipitate) or a "-" (there is none)

2. Obtain an unknown solution which contains two or more ions, but only those used in this experiment. React this unknown with each of the test solutions in sets A and B.

<u>Disposal</u>: Since tiny amounts of chemicals are used, they may be washed down the drain with excess water. Rinse and return your plastic sheets <u>without</u> scrubbing (scratches surface).

Processing:

1. First present your data charts (include all three sets, A, B and C!)

2. Complete the given summary data chart as instructed below. NOTE: The numbers of the chemicals that react to form a precipitate and their formulas are shown in the first column.

In the first two empty columns (NOT the columns that have the numbers!):

a. Write the formulas of the two possible precipitates. These are the products of the double replacement reaction. Be sure to write the correct formulas by considering the two new combinations of ions and balancing their charges.

b. Circle the formula(s) that represents the precipitates. Precipitates will be barium sulfate, barium hydroxide, magnesium hydroxide, barium chromate, silver chromate, or silver chloride.

In the last (long) empty box, write IONIC equations as described here:

c. Write balanced net ionic equations for the different precipitates in sets A and B. Include phase symbols.

d. Write balanced complete ionic equations for all of the precipitate reactions in Set C. Include phase symbols. To do this, on a separate piece of paper, balance the double replacement reaction, then turn everything but the precipitate to ions. Attach the separate sheet to show your work.

3. Logic question: Each unknown (X and Y) contains only a mixture of ions (2 or more) used in sets A and B. For each unknown consider all tests from both sets A and B to make a chart with three columns to list the following:

i. The ions from sets A and B that must be present in the unknown
ii. The ions from sets A and B that may be present in the unknown
iii. The ions from sets A and B that cannot be present from the unknown

Explain briefly how you arrived at your choices:

Hints: 1) Remember, each ion can appear on only one of the three lists for each unknown!
2) Ions involved in precipitates can usually be placed in the "must be" or "cannot be" columns.
3) Spectator ions (those never involved in precipitates) will almost always go in the "may be present" column.

Reactants Set A	Possible (circle the	Precipitates actual ppt)	Reactions Net ionic equation
1. $Ba(NO_3)_2$ 2. Na_2SO_4			
1. $Ba(NO_3)_2$ 3. $Al_2(SO_4)_3$			

Set B			Net ionic equation
1. $MgCl_2$ 2. NaOH			
2. NaOH 4. $BaCl_2$			
2. NaOH 6. $MgSO_4$			
4. $BaCl_2$ 5. Na_2SO_4			
4. $BaCl_2$ 6. $MgSO_4$			

Set C			Complete ionic equation
1. K_2CrO_4 3. $BaCl_2$			
1. K_2CrO_4 5. $AgNO_3$			
2. $Al_2(SO_4)_3$ 3. $BaCl_2$			
3. $BaCl_2$ 4. Na_2CrO_4			
3. $BaCl_2$ 5. $AgNO_3$			
4. Na_2CrO_4 5. $AgNO_3$			